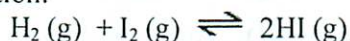


Equilibrium Questions

Consider the following chemical reaction:

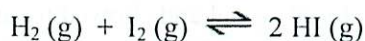


At equilibrium in a particular experiment, the concentrations of H_2 , I_2 , and HI were 0.15 M, 0.033 M, and 0.55 M, respectively.

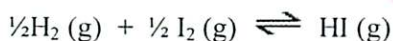
- Write the expression for the equilibrium constant.
- What is value of the equilibrium constant?

$$K_{\text{eq}} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(0.55\text{M})^2}{(0.15\text{M})(0.033\text{M})} = 242.61$$

The value of K_{eq} for the equilibrium

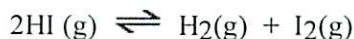


is 794 at 25 °C. What is the value of K_{eq} for the equilibrium below?



$$K_c' = \frac{[\text{HI}]}{[\text{H}_2]^{1/2}[\text{I}_2]^{1/2}} = \sqrt{K} = 28.2$$

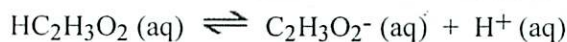
A reaction vessel is charged with hydrogen iodide, which partially decomposes to molecular hydrogen and iodine:



When the system comes to equilibrium at 425 °C, $P_{\text{HI}} = 0.708$ atm, and $P_{\text{H}_2} = P_{\text{I}_2} = 0.0960$ atm.. What is the value of K_p ?

$$K_p = \frac{(P_{\text{H}_2})(P_{\text{I}_2})}{(P_{\text{HI}})^2} = \frac{(0.0960\text{atm})^2}{(0.708\text{atm})^2} = 0.0184$$

Acetic acid is a weak acid that dissociates into the acetate ion and a proton in aqueous solution:



At equilibrium at 25 °C a 0.100 M solution of acetic acid has the following concentrations:

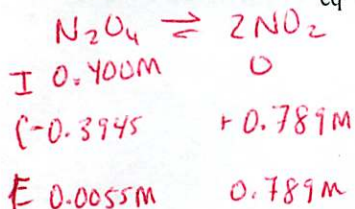
$[\text{HC}_2\text{H}_3\text{O}_2] = 0.0990$ M, $[\text{C}_2\text{H}_3\text{O}_2^-] = 1.33 \times 10^{-3}$ M and $[\text{H}^+] = 1.33 \times 10^{-3}$ M. The equilibrium constant, K_{eq} , for the ionization of acetic acid at 25 °C is _____.

$$K_{\text{eq}} = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = \frac{(1.33 \times 10^{-3}\text{M})^2}{(0.0990\text{M})} = 1.79 \times 10^{-5}$$

Dinitrogen tetroxide partially decomposes according to the following equilibrium:



A 1.00-L flask is charged with 0.400 mol of N_2O_4 . At equilibrium at 373 K, 0.0055 mol of N_2O_4 remains. What is the K_{eq} for this reaction?



$$K_{\text{eq}} = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = \frac{(0.789\text{M})^2}{0.0055\text{M}} = 113$$

At 22 °C, $K_{\text{p}} = 0.070$ for the equilibrium:



A sample of solid NH_4HS is placed in a closed vessel and allowed to equilibrate. Calculate the equilibrium partial pressure (atm) of ammonia, assuming that some solid NH_4HS remains.

$$K_{\text{p}} = [\text{NH}_3][\text{H}_2\text{S}] = 0.070$$

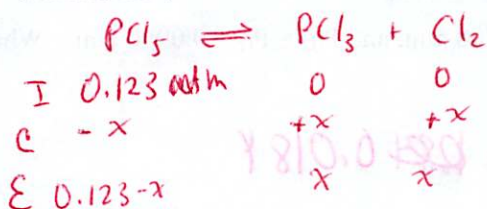
$$x^2 = 0.070$$

$$x = 0.26\text{ atm} = P_{\text{NH}_3}$$

The equilibrium constant (K_{p}) for the interconversion of PCl_5 and PCl_3 is 0.0121:



A vessel is charged with PCl_5 , giving an initial pressure of 0.123 atm. At equilibrium, the partial pressure of PCl_3 is _____ atm.



$$K_{\text{p}} = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]} = 0.0121$$

$$\frac{x^2}{0.123-x} = 0.0121$$

$$x^2 + 0.0121x - 0.0049 = 0$$

$$x = 0.0330\text{ atm} = [\text{PCl}_3]$$

For the reaction below, $K_{\text{p}} = 0.0198$ at 721 K.



In a particular experiment, the partial pressures of H_2 and I_2 at equilibrium are 0.710 and 0.888 atm, respectively. The partial pressure of HI is _____ atm.

$$K_{\text{p}} = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = 0.0198$$

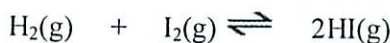
$$\frac{(0.710\text{ atm})(0.888\text{ atm})}{x^2} = 0.0198$$

$$31.8 = x^2$$

$$5.64\text{ atm} = x$$

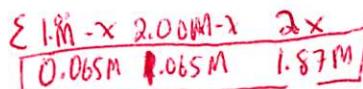
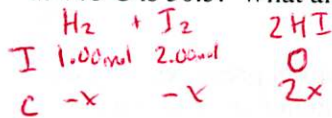
More Complex Equilibrium Problems

A 1.00 L flask is filled with 1.000 mol of H_2 and 2.000 mol of I_2 at $448^\circ C$. The value of the equilibrium constant K_c for the reaction



$$K_c = \frac{[HI]^2}{[H_2][I_2]} = 50.5$$

at $448^\circ C$ is 50.5. What are the equilibrium concentrations of H_2 , I_2 , and HI in moles per liter.



$$\frac{(2x)^2}{(1.00-x)(2.00-x)} = 50.5$$

$$4x^2 = 50.5(x^2 - 3.00x + 2.00)$$

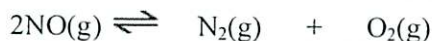
$$46.5x^2 - 151.5x + 101.0 = 0$$

$$\frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$\frac{-(-151.5) \pm \sqrt{(-151.5)^2 - 4(46.5)(101.0)}}{2(46.5)}$$

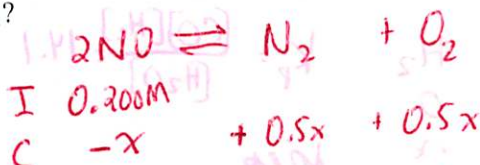
$$= 2.323 \text{ or } 0.935$$

At $2000^\circ C$ the equilibrium constant for the reaction



$$K_c = \frac{[N_2][O_2]}{[NO]^2} = 2.4 \times 10^3$$

is $K_c = 2.4 \times 10^3$. If the initial concentration of NO is 0.200 M, what are the equilibrium concentrations of NO , N_2 , and O_2 ?

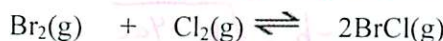


$$\frac{(0.5x)^2}{(0.200-x)^2} = 2.4 \times 10^3$$

$$\frac{0.5x}{0.200-x} = 2.4 \times 10^3$$

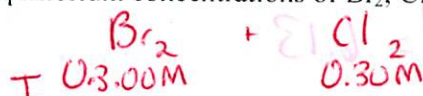
$$0.5x = (2.4 \times 10^3)(0.200-x) \rightarrow x = 0.199$$

For the equilibrium



$$K_c = \frac{[BrCl]^2}{[Br_2][Cl_2]} = 7.0$$

at 400K, $K_c = 7.0$. If 0.30 mol of Br_2 and 0.30 mol of Cl_2 are introduced into a 1.0 L container at 400K, what will be the equilibrium concentrations of Br_2 , Cl_2 , and $BrCl$?



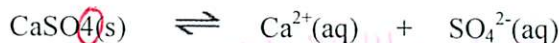
$$K_c = \frac{(2x)^2}{(0.300-x)(0.300-x)} = 7.0$$

$$\frac{2x}{0.300-x} = 2.65$$

$$4.65x = 0.795$$

$$x = 0.171$$

Consider the reaction



$$K_{sp} = [Ca^{2+}][SO_4^{2-}]$$

At $25.0^\circ C$ the equilibrium constant is $K_c = 2.4 \times 10^{-5}$ for this reaction. If excess calcium sulfate is mixed with water at $25.0^\circ C$ to produce a saturated solution, what are the equilibrium concentrations of each ion? If the resulting solution has a volume of 3.0L, what is the minimum mass of calcium sulfate needed to achieve equilibrium?

a $K_{sp} = x^2 = 2.4 \times 10^{-5}$

$$x = 0.0049 M = [Ca^{2+}] = [SO_4^{2-}]$$

b $0.0049 M Ca^{2+} \times 3.0 L = 0.0147 mol Ca^{2+} \times \frac{1 mol CaSO_4}{1 mol Ca^{2+}} \times \frac{136 g CaSO_4}{1 mol CaSO_4} = 2.00 g CaSO_4$

Multiple Concept Problem

At temperatures near 800°C, steam passed over hot coke (a form of carbon obtained from coal) reacts to form CO and H₂:



The mixture of gases that results is an important industrial fuel called *water gas*.

(a) At 800°C the equilibrium constant for the reaction is $K_p = 14.1$. What are the equilibrium partial pressures of H₂O, CO, and H₂ in the equilibrium mixture at this temperature if we start with solid carbon and 0.100 mol of H₂O in a 1.00 L vessel?

(b) What is the minimum amount of carbon required to achieve equilibrium under these conditions?

(c) What is the total pressure in the vessel at equilibrium?

(d) At 25°C the value of K_p for this reaction is 1.7×10^{-21} . Is the reaction endothermic or exothermic?

(e) To produce the maximum amount of CO and H₂ at equilibrium, should the pressure of the system be increased or decreased?

a.

$$P_{\text{H}_2\text{O}} = \frac{nRT}{V} = \frac{(0.100 \text{ mol})(8.314 \text{ J/mol}\cdot\text{K})(1073 \text{ K})}{1.00 \text{ L}} = 8.81 \text{ atm}$$

	C	H ₂ O	CO	H ₂	$K_p = \frac{[\text{CO}][\text{H}_2]}{[\text{H}_2\text{O}]} = 14.1$
I	—	8.81 atm	0	0	
C	—	-x	+x	+x	
E	—	8.81 atm - x	x	x	
L	—	2.67 atm	6.13	6.13	

$$\frac{x^2}{8.81 - x} = 14.1$$

$$x^2 = 124.2 - 14.1x$$

$$x^2 + 14.1x - 124.2 = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-14.1 \pm \sqrt{(14.1)^2 - 4(-124.2)}}{2} = 6.13$$

b 6.13 atm H₂O consumed

$$n = \frac{PV}{RT} = \frac{(6.13 \text{ atm})(0.100 \text{ L})}{(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(1073 \text{ K})} = 0.0696 \text{ mol H}_2\text{O} \times \frac{1 \text{ mol C}}{1 \text{ mol H}_2\text{O}} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.835 \text{ g C}$$

c 2.67 atm + 6.13 atm + 6.13 atm = 14.93 atm

d As $T \downarrow$, $K_p \downarrow$, therefore the Rxn is endothermic b/c lowering T favors the endothermic process. Therefore, this lowers the K_p which means [Reactant] \uparrow and [product] \downarrow . Thus $K_p \downarrow$

e Decreased. As $P \downarrow$ the Equilibrium shifts Right to favor the side of Rxn that has the greater amount of gaseous products.